Nuggets: Quantum: Wavefunction $\Psi$; $\Psi^2$ plots; Radial distribution plot; Nodes; Quantum Numbers: $n$, $l$, $m_l$, $m_s$; Orbital Shapes

Electron Configurations: Orbital Energy Diagram; Orbital Box Diagram; Core vs. Valence Electrons; Shorthand Notation; Charged Elements; Hund’s Rule; Exceptions; Half-filled/Filled Subshells; Charged Transition Metals; #Unpaired Electrons; f-elements; Isoelectronic; Isovalent

Schrodinger’s Equation $\mathcal{H}\Psi = \hat{\mathcal{H}}\Psi$

- Wave function, $\Psi$ (Greek letter psi, pronounced “sigh”) are equations that describe an electron being treated as a stationary wave. The solution to $\Psi(x,y,z)$ are 3 variables: $n$, $l$, and $m_l$

- $\Psi^2(x, y, z)$ – the probability of finding an electron at a given point near position (x, y, z); the plot describes electron density; the plots asymptotically approach the x-axis but do not reach zero probability

- Radial distribution – the total probability of finding an electron at a given distance, $r$, from the nucleus (the equation, $\Psi^2(4\pi r^2)$, is for s subshells); treat atoms as onion layers – as $r$ increases the surface area increases; therefore there are more points at distance $r$ from the nucleus each with their own probability of the electron being there, and the sum of all of these points’ probability is plotted versus $r$; the plots asymptotically approach the x-axis but do not reach zero

- Orbital – a surface containing a region in space in which the electron is found 90% of the time

$\Psi^2$ Probability or Electron Density Plots for 1s, 2s, and 3s orbitals

$\Psi^2$ Radial Distribution/Probability Plots for 1s, 2s, 3s, 2p, and 3p (Radial plots created by José Laboy)
Cross Section and Traditional Boundary Plots for 1s, 2s, and 3s orbitals

Cross section plots for 1s, 2s, and 3s orbitals and their nodes (upper) and the traditional 1s, 2s, and 3s orbitals (lower images).

The boundary/shapes of the s, p and d orbitals

Shown above: one 1s-orbital, one 2s-orbital, three 2p-orbitals, one 3s-orbital, and three 3p-orbitals overlapping. No 3d-orbitals are shown.

Shown above: five 3d-orbitals; these would be overlaid on the orbitals shown to the left (1s, 2s, 2p, 3s and 3p orbitals).
QUANTUM NUMBERS (qn): a set of 3 numbers \((n, l, m_l)\) that is a solution to the Schrödinger wave equation; 
\(n = \text{principal qn}: \text{the energy level/shell/distance from the nucleus}\) of the electron \((n = 1, 2, 3, 4, 5, 6, 7 \ldots)\)
\(l = \text{angular momentum qn or azimuthal qn}: \text{describes orbital shape} (l = 0, 1, \ldots, n-1); l \text{ is the subshell}\)
\(m_l = \text{magnetic qn}: \text{describes orbital orientation} (m_l = -l, \ldots, 0, \ldots, +l)\)
\(m_s = \text{spin qn: } e^+ \text{ spin} \) ("up" or "down"; not part of the Schrödinger equation solution; determined later from the Stern-Gerlach experiment); 
\(m_s = \pm \frac{1}{2}; \text{ magnetic field pointing up or down}\)

- \(n^2 = \#\text{orbitals in the } n^{th}\text{ energy level} \) (\#orbitals in the \(5^{th}\text{ energy level? } n^2 = 5^2 = 25\text{ orbitals})
- \(2n^2 = \#\text{electrons in the } n^{th}\text{ energy level} \) (\#e\(^-\) in the 5\(^{th}\text{ energy level? } 2n^2 = 2(5^2) = 50e^-\))
- \(2l + 1 = \#\text{orbitals in the } l\text{-subshell} \) (\#orbital in the f-subshell? \(2l+1 = 2(3) + 1 = 7\text{ orbitals})
- Each orbital can hold a maximum of 2 electrons
- Pauli Exclusion Principle: No two e\(^-\) can have the same set of 4 qn; there is a unique set of 4 qn for each e\(^-\)
- Each electron defined by a set of 4 qn \((n, l, m_l, m_s)\); each orbital defined by a set of 3 qn \((n, l, m_l)\)

From different \(n \) and \( l \) qn subshells are generated: 1s, 2s, 2p, 3s, 3p, 3d, 4s, 4p, 4d, 4f, 5s, 5p, 5d, 5f, 5g, etc.

\((s = \text{sharp}); \ p = \text{principal}; \ d = \text{diffuse}; \ f = \text{fundamental})

s-subshells contain 1 orbital; p-subshells contain 3 orbitals; d-subshells contain 5 orbitals, etc., (from \(2l + 1\))

Nodes – a point on the previous graphs where the probability of finding an e\(^-\) is zero

Surface Nodes also sometimes called Planar Nodes – are found at the origin in \(\Psi^2\) Probability and \(\Psi^2\) Radial Distribution plots

Radial Nodes – are not at the origin in \(\Psi^2\) Probability and \(\Psi^2\) Radial Distribution plots

\(\#\text{surface/planar nodes} = 1; \#\text{radial nodes} = n - l - 1; \text{ all nodes (surface + radial)} = n - 1;\)

Example 1: How many surface (also called planar) nodes, radial (also called spherical) nodes, and total nodes are there in the 3d subshell?

Answer 1: surface/planar nodes = 2, radial nodes = 0, and total nodes = 2  \{Start by assigning the \(n \) and \( l \) quantum numbers to the 3d subshell: \(n = 3, l = 2\); surface/planar nodes \(l = 2\); radial nodes \(n - l - 1 = 3 - 2 - 1 = 0\); total nodes \(n - 1 = 3 - 1 = 2\)\}

Manipulating Quantum Numbers

Example 2: List all the sets of 4 quantum numbers for the electrons in the 1\(^{st}\) energy level.

Answer 2: \((1, 0, 0, -1/2), (1, 0, 0, +1/2)\)

\(n = 1 \quad \{n = 1 \text{ shell} \to \text{a total of } 2e^- \quad (2n^2 = \text{total } \#e^- \text{ in } n\text{th shell} = 2(1)^2 = 2e^-\}; \text{ therefore, } 2 \text{ sets of } 4 \text{ qn}\)

\(l = 0 \quad \{l \text{ starts at } 0 \text{ and continues until } n-1; \text{ since } n = 1, l \text{ starts at } 0 \to \text{ this is the s-subshell } (l = 0 \to \text{s-subshell}); \text{ there will be } 1 \text{ orbital}
\)

present in the s-subshell \(\#\text{orbitals in s-subshell} = 2l+1 = 2(0) + 1 = 1 \text{ orbital}\); there are a maximum \(2e^-\) per orbital \(\to 2e^- \text{ in s-subshell}\)

\(m_l = 0 \quad \{m_l \text{ starts at } -l, \text{ increments by } 1, \text{ and ends at } +l; \text{ with } l = 0 \to m_l \text{ starts and ends at } 0\}

\(m_s = \pm \frac{1}{2}\)

Sets of quantum numbers are written in the order: \(n, l, m_l, m_s; (1, 0, 0, -1/2), (1, 0, 0, +1/2) \to \text{these two sets of } 4 \text{ qn represent the } 2e^- \text{ in the 1s subshell (subshells are written } \text{“}\text{n-value}\text{” + “l-value”})\)

Example 3: List all the sets of 4 quantum numbers for the electrons in the 2\(^{nd}\) energy level.

Answer 3: \((2, 0, 0, -1/2), (2, 0, 0, +1/2) \) and \((2, 1, -1, -1/2), (2, 1, -1, +1/2), (2, 1, 0, -1/2), (2, 1, 0, +1/2), (2, 1, +1, -1/2), (2, 1, +1, +1/2)\)

\(n = 2 \quad \{n = 2 \text{ shell} \to \text{a total of } 8e^- \quad (2n^2 = \text{total } \#e^- \text{ in } n\text{th shell} = 2(2)^2 = 8e^-\}; \text{ therefore, } 8 \text{ sets of } 4 \text{ qn}\)

\(l = 0 \quad \{l \text{ starts at } 0 \text{ and continues until } n-1; \text{ since } n = 2, l \text{ starts at } 0 \text{ and ends at } 1; \text{ see below for } l = 1; \text{ for } l = 0 \to \text{s-subshell}; \text{ there is } 1 \text{ orbital}
\)

present in the s-subshell \(\#\text{orbitals in s-subshell} = 2l+1 = 2(0) + 1 = 1 \text{ orbital}\); there are a maximum \(2e^-\) per orbital \(\to 2e^- \text{ in s-subshell}\)

\(m_l = 0 \quad \{m_l \text{ starts at } -l, \text{ increments by } 1, \text{ and ends at } +l; \text{ with } l = 0 \to m_l \text{ starts and ends at } 0\}

\(m_s = \pm \frac{1}{2}\)

\((2, 0, 0, -1/2), (2, 0, 0, +1/2) \to \text{these } 2 \text{ sets of quantum numbers represent the } 2e^- \text{ in the 2s subshell}\)

\(l = 1 \quad \{\text{for } l = 1 \to \text{p-subshell}; \text{ there are } 3 \text{ orbitals present in the p-subshell } (2l+1 = 2(1) + 1 = 3 \text{ orbitals}); \text{ since there are a maximum } 2e^- \text{ per orbital } \to 6e^- \text{ in p-subshell}\}

\(m_l = -1, 0, +1 \quad \{m_l \text{ starts at } -l, \text{ increments by } 1, \text{ and ends at } +l; \text{ with } l = 1 \to m_l \text{ starts at } -1 \text{ and increments to } +1\}

\(m_s = \pm \frac{1}{2}\)

\((2, 1, -1, -1/2), (2, 1, -1, +1/2), (2, 1, 0, -1/2), (2, 1, 0, +1/2), (2, 1, +1, -1/2), (2, 1, +1, +1/2) \to \text{these } 6 \text{ sets of quantum numbers represent the } 6e^- \text{ in the 2p subshell}; \text{ Note: } (2s \text{ subshell } = 2e^-) + (2p \text{ subshell } = 6e^-) = 8e^- \text{ (the same total determined from } 2n^2 = (2)^2 = 8e^-)\)
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*Note: Subshells in which no elements occur are not listed; these include: n = 5, l = 4 (5g), n = 6, l = 3 (6h), n = 6, l = 4 (6g); n = 6, l = 5 (6h)

**ELECTRON CONFIGURATIONS**

- **Valence e:** Valence e are the s and p e just past the last noble gas. (Note: Sometimes d-orbital e are also considered valence e; it depends on the specific question.) Valence e are outermost e (those on the atom “surface”); the valence e determine the chemical reactions an atom undergoes

- **Core e:** The e that are not the valence e; the noble gas electrons when using the shorthand notation; the core e are not involved in chemical reactions

- **e fill subshells in a particular order:** 1s2s2p63s23p64s23d104p65s24d105p66s24f145d106p67s25f14
  The superscripts represent the number of electrons within the subshell; the Aufbau diagram outlines the order of filling; (it is better in this author’s opinion to use the Periodic Table to determine e filling order)

- **Shorthand notation:** place prior noble gas symbol in [ ]; this represents all electrons in the noble gas; e.g., Si = 1s2s22p63s23p2 or using the shorthand notation, Si = [Ne]3s23p2 where [Ne] = 1s22s22p6

- **Charged species:** add electrons (anions) or remove electrons (cations) from the last subshell

- **Hund’s rule:** The lowest energy e configuration (called the ground state configuration), is the one with e unpaired as long as possible and the electron spin maximized

- **Charged transition metals: remove s-electrons first when ionizing, not the d-electrons!**

- **Exceptions:** Common ones: Cr, Mo and Cu, Ag, Au (sometimes described: half-filled and filled subshells are more stable than partially filled subshells) (some instructors expect students to have the exceptions memorized)

- **Determine number of unpaired electrons in both neutral and charged species:** ignore electrons in filled subshells and core electrons as these electrons are all paired

- **Diamagnetic:** All e paired (e.g., Mg is diamagnetic)

- **Paramagnetic:** At least 1 e unpaired (e.g., C is paramagnetic; has 2 unpaired electrons)

- **f-Elements (elements past 57): Lanthanides and Actinides; f-electron rows are 4f and 5f (e.g., Pt: [Xe]6s24f145d8)

- **Isoelectronic:** same #e+ (e.g., N2 and CO; #e+ = 14e−);

- **Isovalent:** same #valence e (e.g., Li, Na, and K; ns1; P and O: ns2np4)
In one electron systems, all subshells within one energy level have the same energy. In multi-electron systems, because of e⁻–e⁻ repulsions the subshell energies within one shell split and have different energies (having the same energy is referred to as degenerate). Finally, whichever subshell fills first has lower energy.

Whichever subshell fills first has a lower energy; for example, the 2s subshell is lower in energy than 2p subshell in non-H systems.

Orbital Box Diagram: (Note: The boxes have been placed in order of n value, not the order of filling!)

Example 4: Draw an orbital box diagram for K. Answer 4:

BE ABLE TO WRITE ELECTRON CONFIGURATIONS BY USING THE PERIODIC TABLE

Example 5: What is a possible set of quantum numbers, \( (n, l, m_l, m_s) \), for the last electron in S?

Answer 5: There are six possible answers: \((3, 1, -1, \frac{1}{2})\), \((3, 1, -1, -\frac{1}{2})\), \((3, 1, 0, \frac{1}{2})\), \((3, 1, 0, -\frac{1}{2})\), \((3, 1, 1, \frac{1}{2})\), \((3, 1, 1, -\frac{1}{2})\)

(Write the valence electron configuration for S: \(3\text{s}^23\text{p}^4\); the last electron is placed in the 3p subshell; a 3p subshell has: \(n = 3, l = 1\); if \(l = 1\) then \(m_l = -1, 0, 1\) and \(m_s = \pm \frac{1}{2}\); there are 6 possible sets of 4 quantum numbers \((n, l, m_l, m_s)\) corresponding to the 6 electrons found in the 3p subshell)

Example 6: Write the electronic configuration of Se (shown on Periodic Table above) in the ground state.

Answer 6: Read the PT starting in the upper left corner left to right; when you get to the end of the row on right side, go back all the way to the left side and go down 1 row; as you read across, each element box will represent \(1e^-\); the electrons are written as superscripts; the subshells are shown in the PT above; the PT is organized by electron configurations such that all the subshells are grouped together. The first row: 1s\(^2\); go down 1 row and back to the left for the second row: 2s\(^2\)2p\(^6\); go down 1 row and back to the left for the third row: 3s\(^2\)3p\(^6\); finally go down 1 row and back to the left for the fourth row: 4s\(^2\)3d\(^{10}\)4p\(^6\). Electron configuration for Se: 1s\(^2\)2s\(^2\)2p\(^6\)3s\(^2\)3p\(^6\)4s\(^2\)3d\(^{10}\)4p\(^4\); shorthand notation: [Ar]4s\(^2\)3d\(^{10}\)4p\(^4\)
Example 7: How many unpaired electrons does Mn$^{+3}$ have?

Answer 7: 4 unpaired electrons

Start by writing the electron configuration for Mn$^0$: [Ar]4s$^2$3d$^5$. Change the electron configuration to match the charged metal, Mn$^{+3}$: [Ar]3d$^4$. Note how the first 2 electrons removed are the 4s electrons: With transition metals the s electrons are removed first. The third electron removed is removed from the 3d subshell. Any filled subshell will have zero unpaired electrons so the [Ar] core can be ignored. Draw an orbital energy diagram of the remaining electrons, namely, 3d$^4$:

1. Write all possible sets of 4 quantum numbers for the electrons found in the 3$^{rd}$ energy level. Organize the quantum numbers by subshell (s, p, etc.), label the subshells, and note how many orbitals are in each subshell.

2. Which of the following sets of quantum numbers are not allowed?
   a. n = 4, l = 2, $m_l$ = -1
   b. n = 5, l = 0, $m_l$ = +1
   c. n = 2, l = 1, $m_l$ = 0
   d. n = 3, l = 2, $m_l$ = -2
   e. more than one above

3. Which of the following is not a correct set of quantum numbers for an electron in an atom?
   a. 1 0 0 -1/2
   b. 2 1 -1 +1/2
   c. 3 2 1 -1/2
   d. 4 3 -2 -1/2
   e. All the above sets are correct and possible.

4. For each pair of quantum numbers with the associated atom, do the electrons have the same energy?
   I. For O: (2, 1, 1, $1/2$) and (2, 0, 0, $1/2$)
   II. For N: (2, 1, 1, $1/2$) and (2, 1, -1, -$1/2$)
   III. For Li: (2, 0, 0, $1/2$) and (2, 0, 0, -$1/2$)
   IV. For H: (3, 1, 1, $1/2$) and (3, 0, 0, $1/2$)

5. For each set of conditions, determine how many orbitals will have the listed quantum numbers.
   a. n = 5  b. n = 5, l = 3  c. n = 3, l = 3  d. n = 3, $m_l$ = 2  e. n = 4, l = 2

6. For each set of conditions, determine how many electrons will have the listed quantum numbers.
   a. n = 5  b. n = 5, l = 3  c. n = 4, $m_s = 1/2$  d. n = 2, l = 2
   e. n = 2, $m_s = 0$  f. n = 3, l = 2, $m_l = 1$  g. n = 3, $m_l = 2$

7. Which subshell can have an electron with the magnetic quantum number $m_l$ = -3?
   a. 6s  b. 4f  c. 4d  d. 5p  e. none of these

8. What type of orbital is described by the quantum numbers n = 4, l = 2, $m_l = 0$?
   a. 2s  b. 4f  c. 4d  d. 2g  e. no orbital is described by these numbers

9. I. a. How many surface/planar nodes does the 2p subshell have? b. the 3d subshell?
   II. a. How many radial nodes does the 1s subshell have? b. the 3s subshell?
   III. a. How many total nodes does the 2s subshell have? b. the 4d subshell?

10. Draw the s and p orbitals.
11. a. What is a possible set of four quantum numbers \((n, l, m_l, m_s)\) for the highest energy \(e^-\) in radium (Ra)? (There is more than one possible answer.)
b. What is a possible set of four quantum numbers \((n, l, m_l, m_s)\) for the highest energy \(e^-\) in cesium (Cs)? (There is more than one possible answer.)

12. Write the ground state electron configuration without abbreviations for each atom or ion.
a. Si   b. Cl   c. Fe   d. Kr   e. Ge   f. Sb   g. K+   h. Si\(^{+3}\)   i. S\(^{-2}\)   j. Cu (exception)   k. Cr (exception)

13. Which of the following atoms has the largest number of \textit{valence electrons}?

14. Which of the following is the correct electron configuration for strontium in the ground state?
a. \([\text{Kr}]3s^2\)   b. \([\text{Kr}]4s^2\)   c. \([\text{Kr}]5s^2\)   d. \([\text{Kr}]4d^2\)   e. \([\text{Kr}]4f^2\)

15. For a non-hydrogen atom, which orbital is \textit{lowest in energy}?
a. 2s   b. 4f   c. 4d   d. 2p   e. 3d

16. A series of 9 orbital box diagrams showing the \textit{valence electrons} for C are shown below.
a. Identify which of the diagrams, if any, are \textit{allowed} and are in the \textit{lowest energy state}.
b. Identify which of the diagrams, if any, are \textit{allowed} but are in an \textit{excited state}.
c. Identify which of the diagrams, if any, are \textit{not allowed}.

17. a. Draw an orbital energy diagram for the ground state of N.
b. Draw an orbital energy diagram for an excited state of N. (Note there are an infinite number of possible answers.)

18. a. Write the electron configuration for Mn. Do not use abbreviations.
b. How many unpaired electrons does Mn have?  
c. How many unpaired electrons does Mn\(^{+3}\) have?  
d. List one cation that is isoelectronic with Fe\(^{+2}\)? (There are many possible answers here.)

19. Which atom or ion has the \textit{most unpaired electrons} in the ground state?
a. Si   b. Mn\(^{+2}\)   c. N\(^{-}\)   d. Pb\(^{+2}\)   e. P

20. Which of the following salts will be attracted to a magnet?
I. CuSO\(_4\)   II. Co\(_2\)O\(_3\)   III. Sc(NO\(_3\))\(_3\)   IV. ZnCO\(_3\)

21. In each set, select those elements/ions/compounds which are \textit{isoelectronic}?
a. Cl\(^{-}\), Ar, F\(^{-}\), Br\(^{-}\), Ca\(^+\)   b. CN\(^{-}\), Si, S, Mg\(^{-2}\), Ca\(^{+2}\)
ANSWERS

1. \( n = 3 \) \( \{ n = 3 \rightarrow 18e^- \} \) \( \{ 2n^2 = \text{total #e}^- \text{in n'th shell} = 2(3)^2 = 18e^- \}; \text{therefore, 18 sets of 4 qn} \)
   \( l = 0 \) \{ \( l \) starts at 0 and continues till \( n-1 \); since \( n = 3, l \) starts at 0 and increments by 1 until reaching 2; see below for \( l = 1 \) and \( l = 2 \); \}
   for \( l = 0 \rightarrow s\)-subshell; there is 1 orbital present in \( s\)-subshell \( (2l + 1 = 2(0) + 1 = 1\text{orbital}); \max 2e^-/\text{orbital} \rightarrow 2e^- \text{in s-subshell} \}
   \( m_l = 0 \) \{ \( m_l \) starts at \(-l\), increments by 1, and ends at \(+l\); with \( l = 0 \rightarrow m_l \) starts and ends at 0 \}
   \( m_s = \pm \frac{1}{2} \)
   \( \text{3s subshell:} \ (3, 0, 0, -\frac{1}{2}), (3, 0, 0, +\frac{1}{2}) \rightarrow \text{these 2 sets of qn represent the 2e}^- \text{in the 3s subshell} \)

   \( l = 1 \) \{ \( l = 1 \rightarrow p\)-subshell; there are 3 orbitals present in \( p\)-subshell \( (2l + 1 = 2(1) + 1 = 3\text{orbitals}) \rightarrow 6e^- \text{in p-subshell} \}
   \( m_l = -1, 0, +1 \) \{ \( m_l \) starts at -1 and increments to +1 \}
   \( m_s = \pm \frac{1}{2} \)
   \( \text{3p subshell:} \ (3, 1, -1, -\frac{1}{2}), (3, 1, -1, +\frac{1}{2}), (3, 1, 0, -\frac{1}{2}), (3, 1, 0, +\frac{1}{2}), (3, 1, +1, -\frac{1}{2}), (3, 1, +1, +\frac{1}{2}) \rightarrow \)
   \( \text{these 6 sets of qn represent the 6e}^- \text{in the 3p subshell} \)

   \( l = 2 \) \{ \( l = 2 \rightarrow d\)-subshell; there are 5 orbitals present in \( d\)-subshell \( (2l + 1 = 2(2) + 1 = 5\text{orbitals}) \rightarrow 10e^- \text{in d-subshell} \}
   \( m_l = -2, -1, 0, +1, +2 \) \{ \( m_l \) starts at -2 and increments by 1 to +2 \}
   \( m_s = \pm \frac{1}{2} \)
   \( \text{3d subshell:} \ (3, 2, -2, -\frac{1}{2}), (3, 2, -2, +\frac{1}{2}), (3, 2, -1, -\frac{1}{2}), (3, 2, -1, +\frac{1}{2}), (3, 2, 0, -\frac{1}{2}), (3, 2, 0, +\frac{1}{2}), (3, 2, +1, -\frac{1}{2}), (3, 2, +1, +\frac{1}{2}), (3, 2, +2, -\frac{1}{2}), (3, 2, +2, +\frac{1}{2}) \rightarrow \)
   \( \text{these 10 sets of qn represent the 10e}^- \text{in the 3d subshell} \)

Note: \( (2e^- \text{from 3s subshell}) + (6e^- \text{from 3p subshell}) + (10e^- \text{from 3d subshell}) = 18e^- \) (same total from \( 2n^2 = (2(3)^2 = 18e^- \))

(Note: Students are usually not asked to list all the sets of qn for an entire energy level as this question did; instead this question provides practice in manipulating qn which is a needed skill for other questions often asked.)

2. \( b \) \{if \( l = 0 \), then \( m_l \) must equal 0\}

3. \( e \)

4. \( I. \) no \{the first set of quantum number represents a 2p electron and the second set is for a 2s electron\}
   \( II. \) yes \{the first set of quantum number represents a 2p electron and the second set is also for a 2p electron\}
   \( III. \) yes \{the first set of quantum number represents a 2s electron and the second set is also for a 2s electron\}
   \( IV. \) yes \{the first set of quantum number represents a 3p electron and the second set is for a 3s electron; in one electron systems like H, there is no splitting between subshells; hence, a 2s and 2p electron have equivalent energy; likewise, a 3s, 3p, and 3d electron will also have the same energy, etc.\}

5. \( a. \ 25 \) \{\( n^2 = \#\text{orbitals in n}^{\text{th}} \text{energy level} = 3^2 = 25 \)}
   \( b. \ 7 \) \{\( n = 5 \) and \( l = 3 \rightarrow 5f \text{subshell}; 2l + 1 = 2(3) + 1 = 7 \text{orbitals} \}
   \( c. \ 0 \) \{if \( n = 3 \) then \( l = 0, 1, 2 \) but it cannot equal 3\}
   \( d. \ 1 \) \{if \( n = 3 \) then \( l = 0, 1, \) or 2; if \( l = 0 \) then \( m_l = 0 \) but not 2; if \( l = 1 \) then \( m_l = -1, 0, +1 \) but not 2; if \( l = 2 \) then \( m_l = -2, -1, 0, +1, +2 \)
   \( \rightarrow \text{there is one time } m_l = 2 \text{ which means there is 1 orbital since each } m_l \text{ value corresponds to 1 orbital} \}
   \( e. \ 5 \) \{\( n = 4 \) and \( l = 2 \rightarrow 4d \text{ subshells}; 2l + 1 = 2(2) + 1 = 5 \text{ orbitals} \}
6. a. $50 \{2n^2 = \#e^- \text{ in the } n^{\text{th}} \text{ energy level } = 2(5)^2 = 50\}$
   b. $14 \{n = 5 \text{ and } l = 3 \rightarrow 5f \text{ subshells; } 2l + 1 = 2(3) + 1 = 7 \text{ orbitals } × 2e^- \text{ per orbital } = 14e^-\}$
   c. $16 \{2n^2 = \#e^- \text{ in the } n^{\text{th}} \text{ energy level } = 2(4)^2 = 32e^- \text{ but only } 1/2 \text{ of those } e^- \text{ has } m_s = 1/2 \rightarrow 16e^-\}$
   d. $0 \{\text{if } n = 2 \text{ then } l = 0 \text{ or 1 but it cannot equal 2}\}$
   e. $0 \{\text{no } e^- \text{ can have a } m_s \text{ value equal to 0}\}$
   f. $2 \{\text{when } n = 3 \text{ and } l = 2, \text{ the } m_l \text{ values } = -2, -1, 0, +1, +2; \text{ there is one time } m_l = 1 \text{ which means there is 1 orbital since each } m_l \text{ value corresponds to 1 orbital } \rightarrow 2e^-\}$
   g. $2 \{\text{if } n = 3 \text{ then } l = 0, 1, 2; \text{ if } l = 0 \text{ then } m_l = 0 \text{ but not 2; if } l = 1 \text{ then } m_l = -1, 0, +1 \text{ but not 2; if } l = 2 \text{ then } m_l = -2, -1, 0, +1, +2 \rightarrow \text{ there is one time } m_l = 2 \text{ which means there is 1 orbital since each } m_l \text{ value corresponds to 1 orbital } \rightarrow 2e^-\}$

7. b. $\{\text{for } m_l = 3 \text{ the } l \text{ value must be at least 3 } \rightarrow \text{ it must be at least a } f\text{-subshell; } 4f \text{ in this case}\}$

8. c. $\{l = 2 \rightarrow \text{d-subshell; since } n = 4 \rightarrow 4d \text{ subshell}\}$

9. I. a. 1. $\{\#\text{surface nodes } = l, l = 1 \text{ for } p\text{-subshell } \rightarrow 1 \text{ surface node}\}$
   b. 2. $\{\#\text{surface nodes } = l, l = 2 \text{ for } d\text{-subshell } \rightarrow 2 \text{ surface nodes}\}$
II. a. 0. $\{\#\text{radial nodes } = n - l - 1; \text{ for } 1s \rightarrow n = 1 \text{ and } l = 0 \rightarrow \#\text{radial nodes } = 1-0-1 = 0\}$
   b. 2. $\{\#\text{radial nodes } = n - l - 1; \text{ for } 3s \rightarrow n = 3 \text{ and } l = 0 \rightarrow \#\text{radial nodes } = 3-0-1 = 2\}$
III. a. 1. $\{\#\text{total nodes } = n - 1; \text{ for } 2s \rightarrow n = 2 \rightarrow \text{total nodes } = 2 - 1 = 1\}$
   b. 3. $\{\#\text{total nodes } = n - 1; \text{ for } 4d \rightarrow n = 4 \rightarrow \text{total nodes } = 4 - 1 = 3\}$

10. 

11. a. $(7, 0, 0, +1/2)$ or $(7, 0, 0, -1/2)$  b. $(6, 0, 0, +1/2)$ or $(6, 0, 0, -1/2)$

12. a. $1s^22s^22p^63s^23p^2$
   b. $1s^22s^22p^63s^23p^5$
   c. $1s^22s^22p^63s^23p^64s^23d^6$
   d. $1s^22s^22p^63s^23p^64s^23d104p^6$
   e. $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$
   f. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^3$
   g. $1s^22s^22p^63s^23p^6$
   h. $1s^22s^22p^63s^1$
   i. $1s^22s^22p^63s^23p^6$
   j. $1s^22s^22p^63s^23p^64s^13d^{10}$
   k. $1s^22s^22p^63s^23p^64s^13d^5$

13. e. $\{"a": 3s^23p^3 - 5 \text{ valence } e^-; "b": 3s^23p^1 - 3 \text{ valence } e^-; "c": 5s^2 - 2 \text{ valence } e^-; "d": 2s^22p^3 - 5 \text{ valence } e^-; "e": 2s^22p^4 - 6 \text{ valence } e^-\}$

14. c. $\{\text{Sr: } [\text{Kr}]5s^2\}$
15. a. the subshells filling order is from lowest to highest energy; hence, the 1s subshell is lower in energy than the 2s subshell, etc.; since the 2s subshell from the list of subshells provided will fill first, it is the subshell lowest in energy

16. a. I, II, IV, V {for lowest energy state the electrons should be kept unpaired as long as possible (i.e., placed in separate orbitals) and they should maximize spin (Hund’s rule); to maximize the spin the electrons should point in the same direction as long as possible; up or down makes no difference since electrons don’t know “up from down”}

b. III, VII, VIII {if the electrons are paired prematurely, if the electrons are not pointing in the same direction so the spin is not maximized, or if an electron is placed from a lower energy orbital into a higher energy orbital, then it is an excited state or higher energy state system}

c. VI {no two electrons can have the same four quantum numbers (Pauli Exclusion Principle); in option VI, the 2 electrons are in the same p-orbital pointing in the same direction so each have: n = 2 (since it is a 2p orbital), l = 1 (since it is a p-orbital), m_l = -1 (or 0 or -1; regardless of which m_l value assigned to that orbital the two electrons have the same m_l value), and m_s = +1/2 (since pointing up is usually assigned as a +1/2); hence, both electrons have the same n, l, m_l, m_s values (2, 1, -1, 1/2) which is not allowed}

\[ \frac{1}{2p} \quad \frac{1}{2p} \quad \frac{1}{2p} \]

17. a. \[ \frac{1}{1s} \]

b. Here are 3 possibilities but there are an infinite number:

\[ \frac{1}{1s} \quad \frac{1}{1s} \quad \frac{1}{1s} \]

18. a. 1s^22s^22p^63s^23p^64s^23d^5

b. 5

c. 4

d. Co^{+3} or any cation with 24 electrons

19. b. {Mn: [Ar]4s^23d^5; Mn^{+2}: [Ar]3d^5 since s electrons are removed 1st}

20. b. {determine the charge on the metals; then determine the electron configuration of the metal and see if there are any unpaired electrons present; if there are unpaired electrons then the metal will be attracted to a magnet; recall for charged transition metals to remove s electrons first; CuSO_4: Cu^{+2}: [Ar]3d^6; CoO_2: Co^{+3}: [Ar]3d^6; ZnCO_3: Zn^{+2}: [Ar]3d^{10}; Sc(NO_3)_3: Sc^{+3}: [Ar]3d^0}

21. a. Cl^-, Ar {determine the #e^-: Cl^- = 18e^-; Ar = 18e^-; F^- = 10e^-; Br^- = 36e^-; Ca^+ = 19e^-; find those with the same #e^-}

b. CN^-, Si, Mg^{+2} {determine the #e^-: CN^- = 6 + 7 + 1 = 14e^-; Si = 14e^-; S = 16e^-; Mg^{+2} = 14e^-; Ca^{+2} = 18e^-; find those with the same #e^-}